How's everyone doing?

[STUDENTS CHEERING]

It's a goody bag.

Thank you for that.

It's a goody bag day.

It's a good day.

Related to the goody bags, once again-- hello.

Once again, I want to get these out for you, because you got to have them on a Friday night.

There's no real better way to think about spending your time on a Friday night than opening up a brand new goody bag.

The topic of the goody bag is related to reactions, which is the topic we're starting today.

But it's really also related to acids and bases, which is something we'll talk about on Monday.

So I'm giving you the goody bag today, but I'll talk about the goody bag and contextualize it in the lecture on Monday.

But as I said, today what we're going to do is we're going to start a new topic that we'll be talking about for the next couple of weeks.

And that is the topic of reactions.

And what we're going to talk about is reactions, OK?

Now this falls under the domain of something called chemical kinetics, which is this study of reaction rates.

Kinetic, move, right?

So what's moving?

Concentration, right?

This stuff that you started with is doing something and winding up different.

That's a reaction, OK?

Now we've already written lots of reactions in this class, but today, we're going to talk about the rates.

And the rate in particular is the changing concentration of the reactions and products with time.

So there's a couple of terms that we got to know because we're going to be working with them.

So I'm going to start by writing those down.

So the first thing is the thing that we're going to talk about changing, which is the concentration.

So the concentration is equal to the moles per liter.

I mean, there are lots of ways to write concentration.

This is the way we're going to write it, moles of a substance per liter.

Well, that has a name for it.

That's called them the molarity.

What's molarity?

It's moles per liter.

And then we can also write it in shorthand, which I'll be doing as simply some substance in brackets.

If I say that I have some substance A, and now I say, oh, look, I'm writing brackets around it, then what I mean is the concentration of A, OK?

That's what that means.

Which is the moles per liter of A.

Good, so that's concentration.

What else?

Well, there's the rate.

There is the rate.

And the rate is, as you might guess, the change in the concentration with time.

That's the rate that we're going to be talking about a lot today, the reaction rate.

There's a change in the concentration with time.

What else do we got?

Well, we have then a law.

Because see, here's the thing, the rate is just the rate.

It's the change with something with time, right?

It's like the change in distance with time would be like velocity.

OK, but what if I wanted to say, well, could you tell me what the rate is as some function?

Right, as some function?

That's called the rate law.

That's called the rate law.

And it will give me a rate versus concentration.

That's what I want out of a rate law.

And then, well, I could integrate that.

I could integrate that because the rate itself is a derivative with time.

It's a change in time.

So if I integrate it, well then I've got something called the integrated rate law.

Integrated rate law.

And what's that going to give me?

It's going to give me the concentration versus time.

So the rate law is a rate versus concentration.

And the integrated rate law is the concentration versus time.

And then we've got Arrhenius.

Let's put it all on the same board.

So I'll do one more here, which is the Arrhenius dependence for reactions.

And as you might guess, this is going to give me rate versus temperature.

These are the things we're going to be talking about.

These are all the things-- all of these things we're going to talk about today in the context of reactions.

I wanted to get them all down on the board so you can see them and think about them and feel them as we go forward.

So there are a whole bunch of things that impact the reaction, the rate of a reaction.

What is it that matters?

Well, concentration.

How much do you have?

That's the brackets thing, right?

Temperature, catalysts, the surface and the structure that you're doing the reaction on, the solvent that it's in.

All of these things impact how the reaction happens.

As you can see, I grayed out, made a little darker those bottom two because I'm not going to talk about those two.

But I will talk about each of those other items today.

That's our goal.

How do those items play into these dependencies?

That's our goal, to figure that out.

OK, so we're going to start with concentration.

And we're going to talk about how concentration plays a role in reaction rates.

And we're going to start with something very simple.

I have something called A. Oh, it's purple.

And it turns into something called B. It's green.

There you go.

There's a reaction.

OK, so let's start-- OK, so aA goes to bB.

Now the very first thing that I can think about in this is that nothing is lost.

Nothing is lost.

So I'm going to assume that nothing disappeared from my container.

You see that?

The container is there.

So I haven't lost any mass.

I haven't lost any mass.

Well, mass conservation is a big deal, right?

So mass conservation-- by taking this very simple case, mass conser-- oh, really?

We could have a mass conversation too.

That's OK.

That's kind of what we're doing.

But let's write it as conservation.

This means that if I take away A, I must add B with the same amount and not lose anything.

And if you look at that in terms of these coefficients, remember, these are stoichiometric coefficients as part of the reaction.

Then what that means is that 1 over b times the change in the concentration of B with time must equal minus 1 over a times the change in concentration of A with time.

That is mass conservation, just written as changes in concentration.

You can see it.

Look, imagine just as-- imagine that this is 1, just as an example, OK?

1H2 goes to 2H, right?

I'm not going to lose anything.

Those are my stoichio-- well, OK, this would tell you that the rate of this if I don't lose anything would be 1/2 the change in H, the concentration of H with time equals minus the change in concentration of H2.

That's just exactly what I wrote, but with an example.

I didn't lose anything, right?

So the rate of change of this must be twice the rate of change at that, right?

Stoichiometric coefficients come in and they make sure-- when I write it this way, they make sure I don't lose anything.

That's why you've got to have these there.

Now there's a couple other things about this, all right?

There's a couple other things about this.

One is that you'll notice that the rate is always positive.

Maybe we should write that up on the top board.

Because it's a convention.

But we like positive numbers when we talk about rate.

So the rate is always positive, OK?

Always positive.

So when we have-- when we're thinking about a reaction where something's getting consumed, we're going to write it as the negative of the change in that concentration.

Because that's a negative, and then we're going to write it as a positive of the change in concentration of things being formed, OK?

So the rate is written as positive.

Well, there is another thing, which is we can go more general.

Let's just think about this.

If we go more general, more general, then let's have more than just A to B.

Let's have aA plus bB goes to cC plus dD.

So now I've got four things.

Two products, two reactants, right?

OK A and B are the reactants, C and D are the products.

And they've all got their coefficients.

And so you can look at the change in any one of these and know the rate.

You could look at the change in any one of these and know the rate.

So if I look at this, then the rate would be equal to minus 1 over a, great convention.

OK, it's a positive value.

That's got to be equal to minus 1 over b times the change in B with time.

And it must be equal to plus 1 over C, times the change in C with time.

And that also must be equal to plus 1 over d, all right, times the change in the concentration d change dt.

Right?

So I could look at the way any one of these things is disappearing or forming and because those stoichiometric coefficients are there, I know what the rate of that reaction is.

Again, this is basically a statement of mass conservation.

Ah, but there's more because this is just a definition of rate.

Rate is changing concentration and it's always positive, done.

But I want a law.

I want to know not just OK, yeah, I see how you change with time so that's a rate.

But I want to know a law, that if I gave you any concentration, if I gave you any set of concentrations, what would the rate be?

What would the rate be?

Could I come up with a function?

And so we write down the rate as depending on the concentrations.

Now, you can do this with the products or reactants.

We're going to do this with the reactants.

And so you would have that it depends on the concentrations of a and b raised to some power with some coefficient.

So k times a concentration of A to the m times the concentration of B to the n.

OK, so we're raising these-- Now, we don't know what these are, but what we're saying is that there's got to be a way if I just know, and maybe if I do some experiments or something, maybe I could come up with a general function that depends on just the concentrations, wherever they are of a and b.

And they're going to be raised to some exponent that I don't know yet and they're going to have some constant.

That's a rate law.

That's a rate.

OK?

Now, OK, the thing is that the rate law-- let's go over to here.

OK.

So the k is, as you can imagine going to be very important here, that's called a rate constant.

K is equal to a rate constant.

And this is going to depend on conditions like temperature, pressure, solvent.

So it's going to depend on things like temperature, pressure, solvent, et cetera.

All right?

And we'll see how you get k.

But m and n cannot come-- and this is very important-- m and n and k for that matter must come-- Gesundheit.

--from experiments.

They're determined experimentally, determined experiments.

OK?

This is a mistake that is often made.

You see, m and n must somehow be related to these coefficients.

No.

M and n is something else.

I've created a function that the rate depends on, the rate of this reaction depends on.

I've put these exponents in there and I'm looking for the dependents on m and n.

And the only way to do that, to know that, is to do experiments.

You can't just get that from the way the reaction is written.

You can get mass conservation and the relationship between rate, rate, rate.

Right?

Those are all the same.

But you can't get the rate law unless you do some measurements.

Now, then when you do, then you can get the order of the reaction.

So this is the reaction order and that's an important property.

There's one more thing I'll say, and then what we're going to do, I'm setting the stage here and then we're going to do examples.

OK?

So don't worry, we're going to go through different examples that I think will help crystallize these concepts.

But the last thing I want to say in terms of setting the stage is that the units of rate, if you look at how I've defined the units of rate, it's a change in concentration with time.

So the units of rate, the units of rate are going to be molarity over time.

That doesn't change.

Right?

That doesn't change, those are the units of rate.

That's the definition of the rate.

But you can see, and we will see, that if that has to always be true, then the units of k may very.

And we're going to see that as we go through examples.

So the units of k could be different and we'll be depending on the reaction order, which is basically saying it depends on the rate law.

OK, setting the stage.

Now, let's look at some examples.

And we're going to go through this sequential, we're going to do a zeroth order reaction, than a first order reaction, and then a second order reaction.

OK?

So zeroth order.

Well, zeroth order reaction, right?

So this is nitrous oxide.

So this is a reaction of nitrous oxide turning into N2 and O2.

Nitrous oxide is used in many, many applications, not just laughing gas.

If we go with what we just wrote down.

OK, so let's take this example, I'm going to write it down here.

2 N20 goes to 2 N2 plus O2.

OK.

Well, so the rate we know from what I wrote down, the rate is equal to minus 1/2 times the change in the concentration of N20 with time.

But see, so now I'm doing some measurements and I'm plotting some data here.

So I'm plotting here the concentration of all three of these, the reaction, and the two products with time.

And what you notice is if I plot the concentration with time, it's a linear relationship.

What does that mean?

That means that they don't depend on concentration because the rate is the change right, the rate is the change in those with time, right?

The change in the concentration of time is always the same because it's a straight line.

So if that's true, then the rate from, let's say, straight line, straight line, and we'll talk about plots as we go, straight line plot of concentration versus time.

All right, straight line plot, then we know that concentration is independent of time.

I'm sorry.

Rate is independent.

OK, hold on.

Rate is independent of concentration.

OK.

The rate in the beginning is the change in concentration.

The rate in the middle is the change in concentration with time and it's a straight line so it's always the same.

Well, that means something.

It means something in terms of our rate law.

Because now I know from my rate law, that the rate is equal to k, the rate constant, times N20 to the m.

That's from here, right?

You take all the reactants and you put them down and you put exponents on them.

And so I've just got one, N20 and it's raised to the m.

Yeah, but I know there's no concentration dependents of this rate.

So I know that m equals zero and it's a zeroth order reaction, rxn.

Oh, I've saved so much time, rxn reaction, which means that the rate is simply equal to k.

The rate must equal k, OK?

So in this case the rate, right, so we're continuing here equals k.

And that means that k must have units of molarity over time.

Why, because that's the units of rate, the units of rate don't change, the units of k will.

But in a zeroth order reaction, the units of k have to be equal to the units of rate that there's nothing else in there.

OK.

Again, I want to emphasize, there is no simple correlation between the stoichiometry, the coefficients of these things and the rate law, you've got to do experiments.

This data comes from-- Well look at it, there's the mass conservation in action, 02, N2, all right?

You can see it.

You can feel it.

Slopes are different, they must be.

Now, here's a reaction some of you may care about.

So why does a zeroth order reaction matter.

Why does it matter, because that's the reaction of beer.

So ethanol, OK, but we'll call it beer.

So what does this mean?

If I consume beer, if I did, then the concentration of ethanol in it, there it is plotted versus time is linear.

That means, this is very important, it's a zeroth order reaction.

It means if I plot the rate of the reaction, you see it?

There it is.

It's a constant, it doesn't depend on the concentration.

But that really has a lot of ramifications, because if I drink a lot of beer, the average 70 kilogram person, it takes 2 and 1/2 hours for the enzymes in their liver to decompose 15 milliliters of ethanol, one beer.

But look at that, so that means that if I had more than that the body, it's not dependent on how much I have.

The reaction of the liver, the enzymes that are breaking this down is not dependent on the concentration.

So if I load up a lot more in there it doesn't matter.

There is a pipeline and there's a rate and it's not changing, well that leads to consequences to be thought about when you drink beer.

But there's another thing we can do with beer, which is we can integrate it.

Right?

Because if we know that beer-- if we know that minus-- I'm going to say there's some reaction that beer goes to something.

So minus d beer, dt OK?

Assuming it has a coefficient of 1, is going to equal, that's going to equal k because it's a zeroth order reaction.

Right, it's a constant.

And so OK, now I integrate both sides.

If I integrate, then I've got the integrated right law for beer, which is that the concentration of beer is equal to some initial concentration of beer minus kt.

This is the integrated rate law.

That's the integrated law.

Notice what I've done, I've gone from talking about the rate of the beer being the rate there, which is a constant to a dependence.

There it is, integrated rate law.

The promise was that would get me a concentration versus time and that's what it did, concentration versus time.

How did I get it?

Well, I had a rate law and I integrated it.

And I'm not going through the math, but you put the dt over here and you integrate, OK?

OK.

So that's an integrated rate law.

Well, let's go to the next order.

If I had first order and I looked at the data for that, then you would see if I plot the concentration versus time it would not be a straight line.

Why?

Well, because for a first order reaction, so now we're in first order.

Right, then what that means is that the exponent here is a 1 or if I have two reactants, the addition of them, it's first order total or it might be dependent on two reactants and its first order in each of them, but then second order overall because they're multiplied.

Let's stay we the simpler case of just one reactant A.

So if there's one reactant, and let's suppose A goes to some product.

I'm not even going to write it out.

You know that the rate is equal to minus the change in the concentration of A with time and that because it's first order it depends on the concentration raised to the first power.

That's what we said first order meant, right.

I'm telling you this is first order now.

OK.

Well, you can see that if it's first order it depends on the concentration.

And so that means that you can see from that and you can see from the plot.

If the concentration of A doubles the rate doubles.

The rate doubles.

Why can't our livers do that.

No, they're zeroth order, but there are lots of reactions that are first order.

OK, here's the thing right, but the rate still has to have units of molarity change with time, right molarity over time.

They can't have that unless k units are k units, in this case, are going to be-- So let's see, let me write down-- OK rate has units.

The overall units are molarity over time.
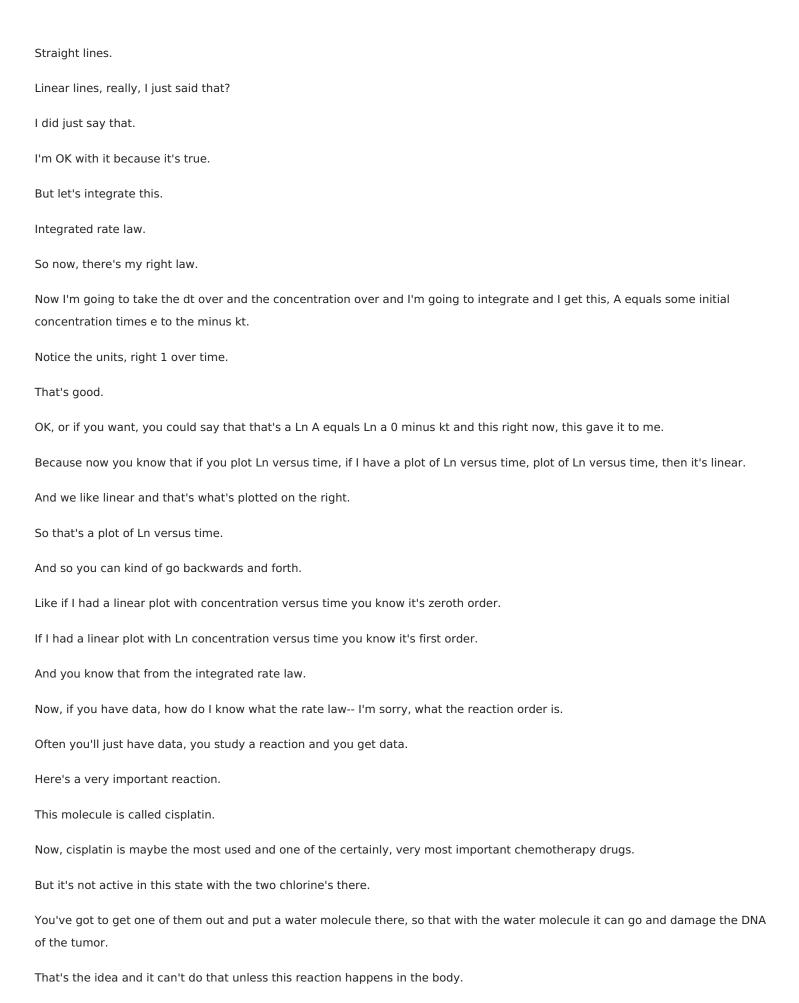
So now I've got k times molarity, that has to have units of molarity over time and so k must have units of 1 over time.

1 over time, but that tells us what the plot.

It helps tell us things about like-- But the plots always come better when we integrate.

If we think about the integrated rate law, then it leads to linearity, which is something we always want in plots.

Linear lines.

Straight lines.

Linear lines, really, I just said that?

I did just say that.

I'm OK with it because it's true.

But let's integrate this.

Integrated rate law.

So now, there's my right law.

Now I'm going to take the dt over and the concentration over and I'm going to integrate and I get this, A equals some initial concentration times e to the minus kt.

Notice the units, right 1 over time.

That's good.

OK, or if you want, you could say that that's a Ln A equals Ln a 0 minus kt and this right now, this gave it to me.

Because now you know that if you plot Ln versus time, if I have a plot of Ln versus time, plot of Ln versus time, then it's linear.

And we like linear and that's what's plotted on the right.

So that's a plot of Ln versus time.

And so you can kind of go backwards and forth.

Like if I had a linear plot with concentration versus time you know it's zeroth order.

If I had a linear plot with Ln concentration versus time you know it's first order.

And you know that from the integrated rate law.

Now, if you have data, how do I know what the rate law-- I'm sorry, what the reaction order is.

Often you'll just have data, you study a reaction and you get data.

Here's a very important reaction.

This molecule is called cisplatin.

Now, cisplatin is maybe the most used and one of the certainly, very most important chemotherapy drugs.

But it's not active in this state with the two chlorine's there.

You've got to get one of them out and put a water molecule there, so that with the water molecule it can go and damage the DNA of the tumor.

That's the idea and it can't do that unless this reaction happens in the body.

And so you can imagine that this kind of reaction is extremely important.

What?

And the rate is extremely important here, all right.

This has to happen in a time frame that we know very well.

So this rate has been studied very carefully for this reaction.

And you can see, OK, you can write down numbers, how much did I start with and how much what was the rate, and I'm watching the rate change.

Right away you know the rate changes as the cisplatin concentration changes, it's not zeroth order.

And the second thing you know is, if you take any two of those and you take ratios, then the ratios are the same so it must be first order.

That's one way to know what order you have.

All right.

Yeah.

So if I have data.

With data like that, with data you could do this.

You could take ratios at two different times.

And what does that get, it gets you the order.

From that you can get the order and we'll see another example.

And then, you can go further than that because then I could plug in data from one of the times, right.

Once I know the order I could use data from one time, from one time to get k.

Because once I know the order, I know how to write down the rate law, and then I can use any of these lines to get k.

So we want to be comfortable with reactions, with rates going back and forth between order, rate law, integrated rate law, plots.

So if I had data that looked like this I could do the same.

Here's another reaction, I could say what order is it?

I'm not going to do it, I'll just say it.

So you take any two times and concentrations and what you have is that the ratio squared gives you the ratio of the rates, that means it must be a second order reaction.

It means it must be, right.

I had something that went into products and the dependence of the rate from this data, the dependence of the rate is that something's concentration squared and you can see it from the data.

So well, let's see.

I'm going back to the middle here.

OK, so now, we just went to second order.

So for second order what do you have?

Well, you have that minus d of something, some reactant, and I just got one reactant here, if you think about it as A. So the same thing A went to something, OK, so minus the change in concentration of that with time is dependent on k times the concentration of that squared.

And if we integrate this one, integrate then what do you get, you get 1 over the concentration of the something equals 1 over its initial value plus kt.

You can see that.

Take the t over here and take this over there and then that's going to be the integrated rate law.

So now, look at that.

What do I plot to get linear?

I got get to linear, I don't like curves in reaction rate plots, I like linear lines because the linear line will tell me something about the reaction order.

And by the way, the slope could tell me something about the rate constant.

Right?

And so if it's second order, the plot that gives you a linear line is 1 over concentration versus time.

So let's look.

So let's say I had this data and I try to plot it.

Well, the first thing I do is I say, well OK hold on, if I plot the concentration versus time that's what it looks like, not linear.

That means it's not zeroth order.

OK, I'm going to take the same data and I'm going to plot the Ln of it versus time, not linear.

Not first order, but if I plot 1 over that concentration versus time it gives me a perfect linear fit, so it must be second order.

Plots, order, linear.

OK?

We're not going to go beyond zeroth, first, and second order because what I want is for you to feel your oneness with reaction rates and kinetics through these three orders, zeroth, 1, and 2, and through these different concepts that are up there.

So one more thing you can do, before I turn to temperature, one more thing you can do once you know the rate law is you can calculate how long it takes for the concentration to be cut in half.

If I know the rate law, then I can calculate the half life, that's called.

All right, so for example, if it's first order let's suppose it's first order then the integrated rate law-- I'm just going to put the two concentrations on the same side here-- A over A0.

OK.

So the integrated rate law gives me this.

It's just what I had before somewhere.

OK, but you can see, well what if the concentration of A is exactly 1/2 of the beginning concentration.

So when A equals 1/2 the initial concentration, but then this becomes a 1/2.

Right?

So if I want to know how much time it takes to get to 1/2, that's a good thing to know just in general.

How long do I have until half of this is left?

Well, then you know that this is going to give you ln2.

When A equals 1/2 A0 then ln2 equals kt.

And we're going to call that an important.

It's not just any time, it's called the half life, it's the time that it took to get to A being 1/2 of the initial.

So I've got it.

I've got it, I've got the half life.

The half life is a very powerful tool.

A number of you may have heard of carbon dating.

This is just a picture of wild animals, but what happens, those wild animals breath in oxygen and other stuff from the air including stuff that contains a spread of carbon atoms and we know carbon has isotopes.

We know about carbons isotopes, we've already talked about them, c12, c13, c14.

C14 has this wonderful thing about it that it radioactivity decays.

And it turns out that there's a certain concentration of that in the atmosphere at any given time and we can go back in history and know how much concentration there was in the atmosphere.

And we know if these animals are breathing it in.

We know that they should have that concentration at A0, except this isn't a concentration this is a number of atoms.

But It's OK, it's still a first order process.

The radioactive decay of c14 is a first order process.

And so, because of that we can date, we can date things using carbon dating very accurately.

This is a really awesome tool that relies simply on the half life, on knowing the half life.

All right, here's a summary that I thought would be useful.

This is it.

This is everything we just talked about before we switch to temperature.

Rate law, zeroth order, first order.

And you're all like, why did I write this all down, it's all on this page.

No, it's good to-- Did somebody say, yeah.

It's good to write down, it helps you think about things.

So here it all is.

There is the zeroth order, first order, second order, OK?

Integrated rate law that comes from that.

The units of the rate constant.

The linear plot to determine the rate constant.

What gives you what the slope of that line is and then what the half life would be if you just follow this for the other two orders.

OK so this is all that we just talked about and the best way to feel your oneness with this is to do some practice problems more than what we've just done.

The next-- really?

The next topic, I'm not going to spend the same amount of time on temperature and catalysts obviously, but I want to mention how these impacts reaction rates.

We're talking about how fast the concentrations change.

And to do that, what we're going to do is we're going to use the collision theory of reactions.

And I've just written it down here what it is.

What is the collision theory of reactions?

What it is, is it says that they occur when particles collide.

Why does a plus b go to c?

Well, because things collided into each other and they got close enough or they maybe had the right orientation so that new bonds could form or others could break.

And the theory also says that not all collisions result in the formation of product and that there are two factors that matter, the energy of the collision, and the orientation of the particles.

I want to highlight those two factors.

So energy of the collision.

We're going to start with the energy of the collision and we're going to talk about that in terms of where you get that energy.

You get it from heat.

You get it from thermal energy, which means these things are vibrating and moving, maybe if it's a gas, they're bouncing off the walls of the container, that's temperature.

How does that matter?

Well, in order to understand how that matters, we have to understand, we have to think about this-- and this is what collision theory for reactions tells us to do is to think about a reaction in terms of an energy landscape and an energy barrier that you have to overcome.

So here I have this Ea, which is the energy that it's going to take.

I need to put that energy in for the collision to be strong enough.

Think about it as like, well I just kind of didn't collide, I had a very low energy, so I'm below Ea and I can't get over this hill.

It really is if you think about it like pushing a car, did I push it hard enough so it can go over the hill, or is it just pushed a little bit, that's the kinetic energy we're talking about, could it get over that hill.

Because what we're seeing is that for this reaction to happen there's some amount of energy that's needed, that's the activation energy, that's the activation energy.

OK.

So what does that mean?

Well, first of all, you should all be feeling something right now.

And you know what I'm talking about, I'm talking about Svante, Svante Arrhenius.

Right, somebody said it.

You were feeling it.

Why, because I'm talking about a process that has a barrier that is thermally activated, a process that has a barrier that is thermally-- Svante Arrhenius, crickets.

Crickets and then intrinsic carriers.

Activated processes, reactions, right, they have a barrier.

They have barrier.

Now here's the thing though, Ea is typically like maybe an electron volt, and we know that KbT is 0.025 eV.

We've talked about this before and how is that possible.

It's the same as what we talked about before because this is an average energy.

If you put the Boltzmann constant in to room temperature, This is at room temperature-- If you put the Boltzmann constant in at room temperature you're going to get that small energy, but we know that thermal energy is a distribution.

This is something we've talked about, so here it is again.

So this is the Boltzmann distribution is what this distribution is called.

And if I have my energy for activating a reaction, my barrier for the reaction to happen, shown there with the vertical line, it means that I have two temperatures.

So yeah, kt may be small, but there's some tails out there where there is enough thermal energy, there's enough kinetic energy for this collision to lead to the reaction.

So those tails grow larger and larger, the amount of molecules that have enough energy grows larger as you increase the temperature.

That is Arrhenius.

The probability for this to happen grows, that is what Arrhenius gives us.

And so here it is mapped.

This is something that is in the textbook that is a nice diagram because it maps the two together.

You've got the reactant going to the products, there's an activated complex, it needs that activation energy and over there we've turned on the side these kinetic energy distribution plots.

So you can see, when do I have enough?

Well, when I'm over that activation energy and there is going to be some fraction of molecules that have enough and that fraction depends on temperature and that is Arrhenius.

And so what we get is that the rate, the rate k, k the rate is equal to some constant A times e to the minus Ea over KbT.

And remember, we use Kb for one collision, for a molecule or an atom.

We would use R or r for a mole.

Here I'm talking about a single event, one molecule colliding with another to make the reaction happen.

Now, this is called the frequency factor and it's a constant.

Constant, constant, and that's something that you could understand how this might depend on a combination of things.

But it's assumed, incorrectly, but it's a good enough approximation for us, but it's assumed to be independent of temperature so we pull it out and call it a frequency factor.

And it has in it information about whether these things really made the right collision or not, that's what A has in it.

This exponent, Arrhenius is telling us this exponent tells me about that fraction, that fraction, but A tells me about orientation.

So here's an example.

Here's N0 plus 03, and on the top line you can see the N0-- right so, the O's are red and N is blue-- and the N0 is coming in, but it's coming in either way whether it's N or whether it's 0, it's coming into the wrong oxygen.

It's hitting that wrong oxygen and so it doesn't do anything because the one in the middle isn't going to be very reactive, not at least at this temperature.

But then you've got a different case where it comes in a little differently, and in fact, the nitrogen is what's coming in and leading the way to that oxygen on the end and that leads to the reaction.

So you can imagine now it's not just how fast they're moving, which is what the exponent tells us, it's more complicated.

Right?

And this is called the frequency factor because it's all these complicated effects rolled into the frequency of the collision being right, not just the temperature.

OK.

And so this gives us the temperature dependence of the rate constant, this gives us the temperature and dependence of the rate constant, which is very important.

This is not time, this is temperature.

All right?

This is a different thing than everything else we talked about here, which involved time dependence, that's temperature dependent.

And finally, the last thing and just to show you this is my very last point, is that just like before, if I wanted to run a reaction and make it go faster, one way, just like with semiconductors is to increase the temperature, but I don't want to run my phone at 600 Kelvin.

So what do we do, we use chemistry.

I may have reactions that I want to run a lot faster.

In fact, I very often do, and what do I do, I use chemistry and those are called catalysts.

What a catalyst does, and this picture shows it in the context of collision theory, I've got these barriers I'm trying to get over, here it's flipped, it doesn't matter.

There's a barrier here, so the reactant is higher than the product, it's OK, there's still a barrier that I have to overcome.

What a catalyst does is without being consumed, what a catalyst does is it's another material in there where it allows that reaction to happen with a lower barrier.

That's what a catalyst is and you better believe this is important.

So a catalyst lowers that without changing the temperature.

And in just 20 seconds, here's my why this matters.

Catalytic converters is one of the most important technologies for pollution that has been invented in over the last 50 years.

It's changed the game in terms of what comes out of that car.

Why does it work?

It wouldn't work, you wouldn't get rid of these things, these toxic things, you wouldn't get rid of them unless you ran it at 1,000 degrees and you're not doing that underneath your car next to the gas tank.

But if you put the right catalyst in, all of those reactions to get rid of those toxic chemicals can happen much more quickly and efficiently.

Right?

OK.

Have a great weekend.

See you guys on Monday.